

Chapter 2-The Chemical Context of Life

Overview:

- Living organisms and the world they live in are subject to the basic laws of physics and chemistry.
- Biology is a multidisciplinary science, drawing on insights from other sciences.
- Life can be organized into a hierarchy of structural levels.
- At each successive level, additional emergent properties appear.

A. Elements and Compounds

Objectives:

1. Distinguish between an element and a compound.
2. Identify the four elements that make up 96% of living matter.
3. Define the term **trace element** and give an example

Notes:

1. Matter consists of chemical elements in pure form and in combinations called compounds.

- Organisms are composed of **matter**.
 - °Matter is anything that takes up space and has mass.
 - °Matter is made up of elements.
- An **element** is a substance that cannot be broken down into other substances by chemical reactions.
 - °There are 92 naturally occurring elements.
 - °Each element has a unique symbol, usually the first one or two letters of the name. Some of the symbols are derived from Latin or German names.
- A **compound** is a substance consisting of two or more elements in a fixed ratio.
 - °Table salt (sodium chloride or NaCl) is a compound with equal numbers of atoms of the elements chlorine and sodium.
 - °While pure sodium is a metal and chlorine is a gas, they combine to form an edible compound. This change in characteristics when elements combine to form a compound is an example of an emergent property.

2. 25 chemical elements are essential to life.

- **About 25 of the 92 natural elements are known to be essential for life.**
 - °Four elements—carbon (C), oxygen (O), hydrogen (H), and nitrogen (N)—make up 96% of living matter.
 - °Most of the remaining 4% of an organism's weight consists of phosphorus (P), sulfur (S), calcium (Ca), and potassium (K).
- **Trace elements** are required by an organism but only in minute quantities.
 - °Some trace elements, like iron (Fe), are required by all organisms.
 - °Other trace elements are required by only some species.
 - For example, a daily intake of 0.15 milligrams of iodine is required for normal activity of the human thyroid gland.

B. Atoms and Molecules

Objectives:

1. Draw and label a simplified model of an atom. Explain how this model simplifies our understanding of atomic structure.
2. Distinguish between each of the following pairs of terms:
 - a. neutron and proton
 - b. atomic number and mass number
 - c. atomic weight and mass number
3. Explain how the atomic number and mass number of an atom can be used to determine the number of neutrons.
4. Explain how two isotopes of an element are similar. Explain how they are different.
5. Describe two biological applications that use radioactive isotopes.
6. Define the terms **energy** and **potential energy**. Explain why electrons in the first electron shell have less potential energy than electrons in higher electron shells.
7. Distinguish among nonpolar covalent, polar covalent and ionic bonds.
8. Explain why strong covalent bonds and weak bonds are both essential in living organisms.
9. Distinguish between hydrogen bonds and Van Der Waals interactions.
10. Give an example that illustrates how a molecule's shape can determine its biological function.
11. Explain what is meant by a chemical equilibrium.

Misconceptions:

1. The simplified models of the atom (Figure 2.4), electron shells (Figure 2.8), and covalent bonding (Figure 2.11) can confuse students who take them too literally. It is important to make sure that students understand that:
 - Atoms do not have defined surfaces.
 - Electrons do not travel in planetary orbits around the nucleus of the atom.
 - Shared electron pairs are not paired spatially in covalent bonds.
 - Electron shells represent energy levels rather than the position of electrons.
2. Students have difficulty fully grasping the concept of energy, and especially the concept of potential energy. Potential energy can be misunderstood as a substance or fuel that is somehow stored in matter. Explain to students that potential energy is associated with an object's ability to move to a lower-energy state, thus releasing some of the potential energy. Return to the concept of potential energy in discussing electron shells, emphasizing that electrons in different electron shells differ in potential energy rather than in position.
3. Students should recognize that weak bonds play important roles in the chemistry of life, despite the transient nature of each individual bond. Page 42 gives the compelling example of the gecko, able to walk on ceilings because of the van der Waals interactions between the ceiling and the hairs on the gecko's toes. Emphasize that strong and weak bonds are both important in the chemistry of life, and ask students to provide examples illustrating this.

Notes:

1. Atomic structure determines the behavior of an element.

- Each element consists of unique atoms.
- An **atom** is the smallest unit of matter that still retains the properties of an element.
 - °Atoms are composed of even smaller parts, called subatomic particles.
 - °Two of these, **neutrons** and **protons**, are packed together to form a dense core, **the atomic nucleus**, at the center of an atom.
 - °**Electrons** can be visualized as forming a cloud of negative charge around the nucleus.
- Each electron has one unit of negative charge.
- Each proton has one unit of positive charge.
- Neutrons are electrically neutral.
- The attractions between the positive charges in the nucleus and the negative charges of the electrons keep the electrons in the vicinity of the nucleus.
- A neutron and a proton are almost identical in mass, about 1.7×10^{-24} gram per particle.
- For convenience, a smaller unit of measure, the **dalton**, is used to measure the mass of subatomic particles, atoms, or molecules.
 - °The mass of a neutron or a proton is close to 1 dalton.
- The mass of an electron is about 1/2000 that of a neutron or proton.
 - °Therefore, we typically ignore the contribution of electrons when determining the total mass of an atom.
- All atoms of a particular element have the same number of protons in their nuclei.
 - °This number of protons is the element's unique atomic number.
 - °The atomic number is written as a subscript before the symbol for the element. For example, ${}^2\text{He}$ means that an atom of helium has 2 protons in its nucleus.
- Unless otherwise indicated, atoms have equal numbers of protons and electrons and, therefore, no net charge.
 - °Therefore, the atomic number tells us the number of protons and the number of electrons that are found in a neutral atom of a specific element.
- The **mass number** is the sum of the number of protons and neutrons in the nucleus of an atom.
 - °Therefore, we can determine the number of neutrons in an atom by subtracting the number of protons (the atomic number) from the mass number.
 - °The mass number is written as a superscript before an element's symbol (for example, ${}^4\text{He}$).
- The **atomic weight** of an atom, a measure of its mass, can be approximated by the mass number.
 - °For example, ${}^4\text{He}$ has a mass number of 4 and an estimated atomic weight of 4 daltons. More precisely, its atomic weight is 4.003 daltons.
- While all atoms of a given element have the same number of protons, they may differ in the number of neutrons.
- Two atoms of the same element that differ in the number of neutrons are called **isotopes**.
- In nature, an element occurs as a mixture of isotopes.
 - °For example, 99% of carbon atoms have 6 neutrons (${}^{12}\text{C}$).
 - °Most of the remaining 1% of carbon atoms have 7 neutrons (${}^{13}\text{C}$) while the rarest carbon isotope, with 8 neutrons, is ${}^{14}\text{C}$.
- Most isotopes are stable; they do not tend to lose particles.

- Both ^{12}C and ^{13}C are stable isotopes.
- The nuclei of some isotopes are unstable and decay spontaneously, emitting particles and energy.
 - ^{14}C is one of these unstable isotopes, or **radioactive isotopes**.
 - When ^{14}C decays, one of its neutrons is converted to a proton and an electron.
 - This converts ^{14}C to ^{14}N , transforming the atom to a different element.
- Radioactive isotopes have many applications in biological research.
 - Radioactive decay rates can be used to date fossils.
 - Radioactive isotopes can be used to trace atoms through metabolic processes.
- Radioactive isotopes are also used to diagnose medical disorders.
 - For example, a known quantity of a substance labeled with a radioactive isotope can be injected into the blood, and its rate of excretion in the urine can be measured.
 - Also, radioactive tracers can be used with imaging instruments to monitor chemical processes in the body.
- While useful in research and medicine, the energy emitted in radioactive decay is hazardous to life.
 - This energy can destroy molecules within living cells.
 - The severity of damage depends on the type and amount of radiation that the organism absorbs.

2. Electron configuration influences the chemical behavior of an atom.

- Simplified models of the atom greatly distort the atom's relative dimensions.
- To gain an accurate perspective of the relative proportions of an atom, if the nucleus was the size of a golf ball, the electrons would be moving about 1 kilometer from the nucleus.
 - Atoms are mostly empty space.
- When two elements interact during a chemical reaction, it is actually their electrons that are involved.
- The nuclei do not come close enough to interact.
- The electrons of an atom vary in the amount of energy they possess.
- **Energy** is the ability to do work.
- **Potential energy** is the energy that matter stores because of its position or location.
 - Water stored behind a dam has potential energy that can be used to do work turning electric generators.
 - Because potential energy has been expended, the water stores less energy at the bottom of the dam than it did in the reservoir.
- Electrons have potential energy because of their position relative to the nucleus.
 - The negatively charged electrons are attracted to the positively charged nucleus.
 - The farther electrons are from the nucleus, the more potential energy they have.
- Changes in an electron's potential energy can only occur in steps of a fixed amount, moving the electron to a fixed location relative to the nucleus.
 - An electron cannot exist between these fixed locations.
- The different states of potential energy that the electrons of an atom can have are called **energy levels** or **electron shells**.
 - The first shell, closest to the nucleus, has the lowest potential energy.
 - Electrons in outer shells have more potential energy.
 - Electrons can change their position only if they absorb or release a quantity of energy that

matches the difference in potential energy between the two levels.

- The chemical behavior of an atom is determined by its electron configuration—the distribution of electrons in its electron shells.
 - °The first 18 elements, including those most important in biological processes, can be arranged in 8 columns and 3 rows.
 - Elements in the same row fill the same shells with electrons.
 - Moving from left to right, each element adds one electron (and proton) from the element before.
- The first electron shell can hold only 2 electrons.
 - °The two electrons of helium fill the first shell.
- Atoms with more than two electrons must place the extra electrons in higher shells.
 - °For example, lithium, with three electrons, has two in the first shell and one in the second shell.
- The second shell can hold up to 8 electrons.
 - °Neon, with 10 total electrons, has two in the first shell and eight in the second, filling both shells.
- The chemical behavior of an atom depends mostly on the number of electrons in its outermost shell, the **valence shell**.
 - °Electrons in the valence shell are known as **valence electrons**.
 - °Lithium has one valence electron; neon has eight.
- Atoms with the same number of valence electrons have similar chemical behaviors.
- An atom with a completed valence shell, like neon, is nonreactive.
- All other atoms are chemically reactive because they have incomplete valence shells.
- The paths of electrons are often portrayed as concentric paths, like planets orbiting the sun.
- In reality, an electron occupies a more complex three-dimensional space, an **orbital**.
- The orbital represents the space in which the electron is found 90% of the time.
 - °Each orbital can hold a maximum of two electrons.
 - °The first shell has room for a single spherical 1s orbital for its pair of electrons.
 - °The second shell can pack pairs of electrons into a spherical 2s orbital and three dumbbell-shaped 2p orbitals.
- The reactivity of atoms arises from the presence of unpaired electrons in one or more orbitals of their valence shells.
 - °Electrons occupy separate orbitals within the valence shell until forced to share orbitals.
 - The four valence electrons of carbon each occupy separate orbitals, but the five valence electrons of nitrogen are distributed into three unshared orbitals and one shared orbital.
- When atoms interact to complete their valence shells, it is the *unpaired* electrons that are involved.

3. Atoms combine by chemical bonding to form molecules.

- Atoms with incomplete valence shells can interact with each other by sharing or transferring valence electrons.
- These interactions typically result in the atoms remaining close together, held by attractions called **chemical bonds**.
 - °The strongest chemical bonds are covalent bonds and ionic bonds.
- A **covalent bond** is formed by the sharing of a pair of valence electrons by two atoms.
 - °If two atoms come close enough that their unshared orbitals overlap, they will share their

newly paired electrons. Each atom can count both electrons toward its goal of filling the valence shell.

°For example, if two hydrogen atoms come close enough that their 1s orbitals overlap, then they can share a pair of electrons, with each atom contributing one.

- Two or more atoms held together by covalent bonds constitute a **molecule**.
- We can abbreviate the structure of the molecule by substituting a line for each pair of shared electrons, drawing the **structural formula**.
 - °H—H is the structural formula for the covalent bond between two hydrogen atoms.
- The **molecular formula** indicates the number and types of atoms present in a single molecule.
 - °H₂ is the molecular formula for hydrogen gas.
- Oxygen needs to add 2 electrons to the 6 already present to complete its valence shell.
 - °Two oxygen atoms can form a molecule by sharing *two* pairs of valence electrons.
 - °These atoms have formed a **double covalent bond**.
- Every atom has a characteristic total number of covalent bonds that it can form, equal to the number of unpaired electrons in the outermost shell. This bonding capacity is called the atom's **valence**.
 - °The valence of hydrogen is 1.
 - °Oxygen is 2.
 - °Nitrogen is 3.
 - °Carbon is 4.
 - °Phosphorus should have a valence of 3, based on its three unpaired electrons, but in biological molecules it generally has a valence of 5, forming three single covalent bonds and one double bond.
- Covalent bonds can form between atoms of the same element or atoms of different elements.
 - °While both types are molecules, the latter are also compounds.
 - °Water, H₂O, is a compound in which two hydrogen atoms form single covalent bonds with an oxygen atom.
 - This satisfies the valences of both elements.
 - Methane, CH₄, satisfies the valences of both C and H.
- The attraction of an atom for the shared electrons of a covalent bond is called its **electronegativity**.
 - °Strongly electronegative atoms attempt to pull the shared electrons toward themselves.
- If electrons in a covalent bond are shared equally, then this is a **nonpolar covalent bond**.
 - °A covalent bond between two atoms of the same element is always nonpolar.
 - °A covalent bond between atoms that have similar electronegativities is also nonpolar.
 - Because carbon and hydrogen do not differ greatly in electronegativities, the bonds of CH₄ are nonpolar.
- When two atoms that differ in electronegativity bond, they do not share the electron pair equally and form a **polar covalent bond**.
 - °The bonds between oxygen and hydrogen in water are polar covalent because oxygen has a much higher electronegativity than does hydrogen.
 - °Compounds with a polar covalent bond have regions of partial negative charge near the strongly electronegative atom and regions of partial positive charge near the weakly electronegative atom.
- An **ionic bond** can form if two atoms are so unequal in their attraction for valence electrons that one atom strips an electron completely from the other.

°For example, sodium, with one valence electron in its third shell, transfers this electron to chlorine, with 7 valence electrons in its third shell.

°Now, sodium has a full valence shell (the second) and chlorine has a full valence shell (the third).

- After the transfer, both atoms are no longer neutral, but have charges and are called **ions**.
- Sodium has one more proton than electrons and has a net positive charge.
 - °Atoms with positive charges are **cations**.
- Chlorine has one more electron than protons and has a net negative charge.
 - °Atoms with negative charges are **anions**.
- Because of differences in charge, cations and anions are attracted to each other to form an **ionic bond**.
 - °Atoms in an ionic bond need not have acquired their charges by transferring electrons with each other.
- Compounds formed by ionic bonds are **ionic compounds**, or **salts**. An example is NaCl, or table salt.
 - °The formula for an ionic compound indicates the ratio of elements in a crystal of that salt. NaCl is not a molecule, but a salt crystal with equal numbers of Na^+ and Cl^- ions.
- Ionic compounds can have ratios of elements different from 1:1.
 - °For example, the ionic compound magnesium chloride (MgCl_2) has 2 chloride atoms per magnesium atom.
 - Magnesium needs to lose 2 electrons to drop to a full outer shell; each chlorine atom needs to gain 1.
- Entire molecules that have full electrical charges are also called ions.
 - °In the salt ammonium chloride (NH_4Cl), the anion is Cl^- and the cation is NH_4^+ .
- The strength of ionic bonds depends on environmental conditions, such as moisture.
- Water can dissolve salts by reducing the attraction between the salt's anions and cations.

4. Weak chemical bonds play important roles in the chemistry of life.

- Within a cell, weak, brief bonds between molecules are important to a variety of processes.
 - °For example, signal molecules from one neuron use weak bonds to bind briefly to receptor molecules on the surface of a receiving neuron.
 - °This triggers a response by the recipient.
- Weak interactions include ionic bonds (weak in water), hydrogen bonds, and van der Waals interactions.
- **Hydrogen bonds** form when a hydrogen atom already covalently bonded to a strongly electronegative atom is attracted to another strongly electronegative atom.
 - °These strongly electronegative atoms are typically nitrogen or oxygen.
 - °These bonds form because a polar covalent bond leaves the hydrogen atom with a partial positive charge and the other atom with a partial negative charge.
 - °The partially positive-charged hydrogen atom is attracted to regions of full or partial negative charge on molecules, atoms, or even regions of the same large molecule.
- For example, ammonia molecules and water molecules interact with weak hydrogen bonds.
 - °In the ammonia molecule, the hydrogen atoms have partial positive charges, and the more electronegative nitrogen atom has a partial negative charge.
 - °In the water molecule, the hydrogen atoms also have partial positive charges, and the oxygen atom has a partial negative charge.
 - °Areas with opposite charges are attracted.

- Even molecules with nonpolar covalent bonds can have temporary regions of partial negative and positive charge.
 - °Because electrons are constantly in motion, there can be periods when they accumulate by chance in one area of a molecule.
 - °This creates ever-changing regions of partial negative and positive charge within a molecule.
- Molecules or atoms in close proximity can be attracted by these fleeting charge differences, creating **van der Waals interactions**.
- While individual bonds (ionic, hydrogen, van der Waals) are weak and temporary, collectively they are strong and play important biological roles.

5. A molecule's biological function is related to its shape.

- The three-dimensional shape of a molecule is an important determinant of its function in a cell.
- A molecule with two atoms is always linear.
- However, a molecule with more than two atoms has a more complex shape.
- The shape of a molecule is determined by the positions of the electron orbitals that are shared by the atoms involved in the bond.
 - °When covalent bonds form, the orbitals in the valence shell of each atom rearrange.
- For atoms with electrons in both *s* and *p* orbitals, the formation of a covalent bonds leads to hybridization of the orbitals to four new orbitals in a tetrahedral shape.
- In a water molecule, two of oxygen's four hybrid orbitals are shared with hydrogen atoms. The water molecule is shaped like a V, with its two covalent bonds spread apart at an angle of 104.5°.
- In a methane molecule (CH₄), the carbon atom shares all four of its hybrid orbitals with H atoms. The carbon nucleus is at the center of the tetrahedron, with hydrogen nuclei at the four corners.
- Large organic molecules contain many carbon atoms. In these molecules, the tetrahedral shape of carbon bonded to four other atoms is often a repeating motif.
- Biological molecules recognize and interact with one another with a specificity based on molecular shape.
- For example, signal molecules from a transmitting cell have specific shapes that bind to complementary receptor molecules on the surface of the receiving cell.
 - °The temporary attachment of the receptor and signal molecule stimulates activity in the receptor cell.
- Molecules with similar shapes can have similar biological effects.
 - °For example, morphine, heroin, and other opiate drugs are similar enough in shape that they can bind to the same receptors as natural signal molecules called endorphins.
 - °Binding of endorphins to receptors on brain cells produces euphoria and relieves pain. Opiates mimic these natural endorphin effects.

6. Chemical reactions form and break chemical bonds.

- In **chemical reactions**, chemical bonds are broken and reformed, leading to new arrangements of atoms.

- The starting molecules in the process are called **reactants**, and the final molecules are called **products**.
- In a chemical reaction, all of the atoms in the reactants must be present in the products.
 - °The reactions must be “balanced.”
 - °Matter is conserved in a chemical reaction.
 - °Chemical reactions rearrange matter; they do not create or destroy matter.
- For example, we can recombine the covalent bonds of H_2 and O_2 to form the new bonds of H_2O .
- In this reaction, two molecules of H_2 combine with one molecule of O_2 to form two molecules of H_2O .
- Photosynthesis is an important chemical reaction.
 - °Humans and other animals ultimately depend on photosynthesis for food and oxygen.
 - °Green plants combine carbon dioxide (CO_2) from the air and water (H_2O) from the soil to create sugar molecules and release molecular oxygen (O_2) as a by-product.
 - °This chemical reaction is powered by sunlight.
 - °The overall process of photosynthesis is $6\text{CO}_2 + 6\text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2$.
 - °This process occurs in a sequence of individual chemical reactions that rearrange the atoms of the reactants to form the products.
- Some chemical reactions go to completion; that is, all the reactants are converted to products.
- Most chemical reactions are reversible, with the products in the forward reaction becoming the reactants for the reverse reaction.
- For example in this reaction: $3\text{H}_2 + \text{N}_2 \rightleftharpoons 2\text{NH}_3$ hydrogen and nitrogen molecules combine to form ammonia, but ammonia can decompose to hydrogen and nitrogen molecules.
 - °Initially, when reactant concentrations are high, they frequently collide to create products.
 - °As products accumulate, they collide to reform reactants.
- Eventually, the rate of formation of products is the same as the rate of breakdown of products (formation of reactants), and the system is at **chemical equilibrium**.
 - °At equilibrium, products and reactants are continually being formed, but there is no net change in the concentrations of reactants and products.

At equilibrium, the concentrations of reactants and products are typically not equal, but their concentrations have stabilized at a particular ratio.

Key Terms

anion

atom

atomic mass

atomic nucleus	hydrogen bond	potential energy
atomic number	ion	product
cation	ionic bond	proton
chemical bond	ionic compound	radioactive isotope
chemical equilibrium	isotope	reactant
chemical reaction	mass number	salt
compound	matter	structural formula
covalent bond	molecular formula	trace element
dalton	molecule	valence
electron	neutron	valence electron
electron shell	nonpolar covalent bond	valence shell
electronegativity	orbital	van der Waals
element	periodic table of the	interactions
energy	elements	
energy level	polar covalent bond	

Word Roots

an- = not (*anion*: a negatively charged ion)

co- = together; **-valent** = strength (*covalent bond*: an attraction between atoms that share one or more pairs of outer-shell electrons)

electro- = electricity (*electronegativity*: the tendency for an atom to pull electrons towards itself)

iso- = equal (*isotope*: an element having the same number of protons and electrons but a different number of neutrons)

neutr- = neither (*neutron*: a subatomic particle with a neutral electrical charge)

pro- = before (*proton*: a subatomic particle with a single positive electrical charge)